Chemistry – Unit 7 Objectives
Stoichiometry I

By the time we finish this unit, you should be able to these:

1. Review Concepts: a) Determine the molar mass of a substance and use it to convert between the mass and mole measurements. (U5) b) Relate coefficients and formulas to a molecular diagram of a reaction. (U6) c) Given a chemical reaction stated in words, write a balanced chemical equation. (U6)	
2. Starting with • a balanced chemical equation, • the number of **moles** of a reactant or product, determine the number of **moles** of any other reactant or product involved.	
3. Starting with • a balanced chemical equation, • the **mass** of a reactant or product, determine the **mass** of any other reactant or product involved.	

4. Starting with • a balanced chemical equation, • the **mass** of one reactant, • **mass** of product **actually** produced calculate the **percent yield** for the reaction	
5. Starting with • a balanced chemical equation, • the **mass** of the reactants. determine • which reactant is **limiting**, and why it limits the reaction, • the **theoretical yield** of a product.	
6. Given a balanced chemical equation and the amounts of reactants, sketch molecular diagrams to represent the reaction mixture before and after the reaction.	
Vocabulary to understand, distinguish, and use correctly: • Stoichiometry • Stoichiometric mole ratio • Theoretical yield • Actual yield • Percent yield • Limiting reactant	

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Chemistry Unit 7
More practice in writing and balancing equations

Write balanced chemical equations for the following reactions.

1. Zinc reacts with hydrogen chloride to form zinc chloride.

2. Sodium oxide reacts with water to form sodium hydroxide.

3. Iron metal reacts with water to form Fe_3O_4 and hydrogen gas.

4. Aluminum bromide reacts with chlorine gas to produce aluminum chloride and liquid bromine.

5. Nitric acid (HNO_3) reacts with barium hydroxide to produce barium nitrate and water.

6. Calcium sulfite decomposes when heated to form calcium oxide and sulfur dioxide.

7. Iron reacts with sulfuric acid (H_2SO_4) to form iron(II) sulfate and hydrogen gas.

8. Ammonia (NH_3) burns in air to form nitrogen dioxide and water.

9. Carbon disulfide burns in air to form carbon dioxide and sulfur dioxide.

10. Lead (II) carbonate reacts with nitric acid to form lead (II) nitrate, water and carbon dioxide.

11. Hydrogen peroxide (H_2O_2) decomposes to produce liquid water and oxygen gas.

12. Magnesium chlorate decomposes to form magnesium chloride and oxygen gas.

Sodium sulfide reacts with nickel(II) nitrate to form nickel sulfide and sodium nitrate.

14. Aluminum, when hot enough, burns in air to form aluminum oxide.

15. Copper(II) oxide, heated in the presence of methane gas (CH_4), produces pure copper metal and the gases carbon dioxide and water.

16. Solutions of sodium carbonate and iron(III) chloride react to form solid iron(III) carbonate and sodium chloride in solution.

Name:________________________________

Unit 7 Worksheet 1: Mole relationships

For each of the problems below:

a. Write the balanced chemical equation

b. Identify what is given (with units) and what you want to find (with units)

c. Use coefficients from balanced equation to determine mole ratio.

d. Show set up (organize it!).

1. Hydrogen sulfide gas, which smells like rotten eggs, burns in air to produce sulfur dioxide and water. How many moles of oxygen gas would be needed to completely burn 8 moles of hydrogen sulfide?

Equation: ___ $H_2S_{(g)}$ + ___ $O_{2\,(g)}$ → ___ $SO_{2(g)}$ + ___ $H_2O_{(g)}$

Before: ___ ___ ___ ___

Change ___ ___ ___ ___

After ___ ___ ___ ___

2. Propane, C_3H_8, burns in air to form carbon dioxide and water. If 12 moles of carbon dioxide are formed, how many moles of propane were burned?

Equation:

Before:

Change

After

3. Ammonia, NH_3, for fertilizer is made by causing hydrogen and nitrogen to react at high temperature and pressure. How many moles of ammonia can be made from 0.15 moles of nitrogen gas?

Equation:

Before:

Change

After

4. The poison gas phosgene, $COCl_2$, reacts with water in the lungs to form hydrochloric acid and carbon dioxide. How many moles of hydrochloric acid would be formed by 0.835 moles of phosgene?

   Equation:

   Before:

   Change

   After

5. Iron metal and oxygen combine to form the magnetic oxide of iron, Fe_3O_4. How many moles of iron can be converted to magnetite by 8.80 moles of pure oxygen? (make your BCA table)

   How many moles of iron oxide would be produced?

6. The recipe for Coca-Cola Classic is a closely guarded secret. Researchers outside the company believe the flavoring mixture, known as "7X", contains oils of orange, lemon, nutmeg, cinnamon, and coriander. The original mixture also contained caffeine, vanilla, caramel, lime juice, sugar or artificial sweetener, and citric acid.
   Over the years, the recipe has changed. For example, the original recipe contained citric acid but this was combined with phosphoric acid to cut production costs. Corn syrup replaced sugar for the same reason.

   $$\underbrace{C_8H_{10}N_4O_2}_{\text{caffeine}} + \underbrace{4\ H_3PO_4}_{\text{phosphoric acid}} + 6\ CO_2 + \text{other ingredients} \rightarrow \underbrace{C_6H_5CO_2K}_{\text{potassium benzoate}} + \text{other products}$$

   To produce 1000 cans of Coca-Cola Classic, 40g (0.21 moles) of caffeine are reacted with phosphoric acid and other ingredients. How many moles of phosphoric acid are required? How many moles of carbon dioxide are required?

Stoichiometry: Predicting Amounts in Reactions

Stoichiometry is the process of determining how much product is made or how much reactant is needed during a chemical reaction. As we know, in chemical reactions atoms are **conserved**. We show this in a balanced chemical equation.

The balanced chemical equation tells us two things:

1. **Which substances** begin with (reactants) and end with (products) during the rearrangement process.
2. The **ratio of particles** involved. This ratio can be seen either as a ratio of individual particles OR as a ratio of moles.

In the lab it is only practical to work with **moles** of substances rather than individual atoms or molecules, and so we interpret our equations as a ratio of moles, or a **mole ratio.**

Example: $2\ Mg \quad + \quad 1O_2 \quad \rightarrow \quad 2\ MgO$ means
for every *2 moles of Mg* burned, *1 mole of O_2* is required to produce *2 moles of MgO*, or a ratio of ***2 moles Mg : 1 mole O_2 : 2 moles MgO***

We can use this mole ratio relationship to make predictions about how much we need of something, or how much we can make from what we have.

Making Predictions
In every reaction, there are three stages we need to consider to make good predictions:

1. **Before**: What we have before the reaction takes place.
2. **Change**: How much of each substance *actually changes* during the reaction
3. **After**: How much of each substance is present after the reaction is complete.

Some good organization can help us in making good predictions. We have an organizational table that can help us track the Before-Change-After for a reaction. Below is an example of a problem involving a chemical reaction.

Sample Problem: Hydrogen sulfide gas, which smells like rotten eggs, burns in air to produce sulfur dioxide and water. How many moles of oxygen gas would be needed to completely burn 2.4 moles of hydrogen sulfide?

Step 1- Write and Balance the equation (describe the reaction and its mole ratio)

$$2\ H_2S \ + 3\ O_2 \rightarrow 2\ SO_2 \ + \ 2\ H_2O$$

Before:

Change

After

Step 2: Fill in the *Before* line with the <u>Given</u> information; mark what you must <u>Find</u> on the table (with units)

$2\ H_2S + 3\ O_2 \rightarrow 2\ SO_2 + 2\ H_2O$

	$2\ H_2S$	$3\ O_2$	$2\ SO_2$	$2\ H_2O$
Before:	***2.4moles***	***xs moles***	***0 moles***	***0 moles***
Change		***__ moles***		
After				

NOTE: Assume reactants you don't have amounts for are present with <u>more than enough</u> available (**excess**, or "**xs**") for the reaction to be completed.

Step 3: Use ratio of coefficients to determine the *Change* made

$\mathbf{2}\ H_2S + \mathbf{3}\ O_2 \rightarrow \mathbf{2}\ SO_2 + \mathbf{2}\ H_2O$

	$2\ H_2S$	$3\ O_2$	$2\ SO_2$	$2\ H_2O$
Before:	2.4moles	xs moles	0 moles	0 moles
Change:	***–2.4 moles***	***–<u>3.6</u> moles***	***+2.4 moles***	***+2.4 moles***
After				

NOTE: Reactants are consumed/decrease (-), products accumulate/increase (+)

Step 4: Complete the table for what remains *After* the reaction is complete

$2\ H_2S + 3\ O_2 \rightarrow 2\ SO_2 + 2\ H_2O$

	$2\ H_2S$	$3\ O_2$	$2\ SO_2$	$2\ H_2O$
Before:	2.4 moles	xs moles	0 moles	0 moles
Change:	–2.4 moles	–<u>3.6</u> moles	+2.4 moles	+2.4 moles
After:	***0 moles***	***xs moles***	***2.4 moles***	***2.4 moles***

- In this case, desired answer is in moles
  <u>Answer</u> (in moles): <u>3.6 moles O_2</u> are needed to burn 2.4 moles H_2S.

- If mass is required, convert moles to grams in the usual way
  $3.6 \text{moles } O_2 * \frac{32 \text{ grams}}{1 \text{ mole}} = 115 \text{ grams } O_2$
  <u>Answer</u> (in grams): <u>115 grams O_2</u> are needed to burn 2.4 moles H_2S.

Chemistry Unit 7 Lab
Copper-Silver Nitrate Reaction

Introduction

In this experiment, a solution of silver nitrate will react with copper wire. Silver metal will be produced. Careful measurements will enable you to determine the mole relationships between the reactants and products.

Procedure

1. Obtain a length of copper wire, a vial of silver nitrate, and a clean, dry beaker. Label the beaker with the period and group number so you can find it again.

2. Find the mass of each of these to the nearest 0.01g. Use the same balance for all your measurements in this experiment. To be sure, record the balance number.

3. Clamp a clean 18 x 150mm test tube to a ring stand. Add about 20 mL of **distilled** water to the test tube. Add the silver nitrate from the vial to the test tube, stir gently to dissolve the crystals. CAUTION: silver nitrate, solid or solution, will stain your skin and clothing. Avoid contact, and rinse immediately with water if contact occurs. Be sure to rinse the stirring rod with a small amount of distilled water into the test tube when you are done stirring the solution.

4. Coil the copper wire by wrapping it around a pencil. Stretch the wire until it is about 2 cm longer than the test tube. Place the coiled wire into the test tube with the silver nitrate solution so that the straightened end is out of the solution. Note the reaction that occurs; record your observations. Allow the reaction to continue for 30 minutes. While the reaction is taking place, find the mass of the empty silver nitrate vial to the nearest 0.01g.

5. Shake the silver crystals from the copper wire and remove the wire from the test tube. Using the wash bottle, rinse the wire into the weighed, labeled beaker. Pour the contents of the test tube into the beaker (you may need to use distilled water wash bottle to rinse it out). Dip the copper wire in acetone, then set it aside to dry. When it is dry, find its mass to the nearest 0.01g, then set it in the place directed by your instructor.

6. Carefully decant the solution from the silver crystals into another beaker. Rinse the silver with about 10 mL of distilled water, stirring gently with the rod. Allow the silver to settle, then decant into the other beaker. Repeat this 3 more times. Don't worry about the few tiny particles of silver that float off as you decant. Their mass is negligible. Discard the liquid from the other beaker.

7. After the final rinse, place the beaker with silver in the drying oven.

8. When the silver is dry, find the mass of the beaker and silver to the nearest 0.01g. Then, place the beaker in the fume hood, and pour 5.0 mL of 6M HNO_3 (nitric acid) into the beaker. The nitric acid reacts with the silver producing the reddish-brown NO_2 gas (quite noxious, do not smell it) and turning the silver back into silver nitrate, $AgNO_3$. This will be used in the next experiment.

Data

balance # _________

mass of beaker ____________g

mass of vial + $AgNO_3$ ____________g

mass of Cu wire before ____________g

mass of empty vial ____________g

mass of Cu wire after ____________g

mass of beaker + dry Ag ____________g

Calculations

1. Determine the mass of copper that reacted during the experiment. Convert this to moles of Cu.
2. Determine the mass of silver produced during the experiment. Convert this to moles of Ag.
3. Determine the value of the ratio: $\frac{moles Ag}{moles Cu}$. Be sure to use the appropriate SF.
4. Determine the mass and then the number of moles of $AgNO_3$ used in the lab.
5. Determine the value of the ratio: $\frac{moles\ Ag}{moles\ AgNO_3}$. Be sure to use the appropriate SF.

Conclusion

1. Because of our deep and abiding belief that atoms react in simple integer ratios, what do you suppose is the *actual* ratio of $\frac{moles\ Ag\ \text{produced}}{moles\ Cu\ \text{consumed}}$?

   You should use class results to help you decide the correct value.
   From your answer, write the balanced equation for the reaction between copper and silver nitrate.

2. Make a BCA table for this chemical reaction Enter the value of moles $AgNO_3$ in the before line and xs for moles of Cu. In the change line, enter the moles of Cu reacted. Now, use the moles of Cu to complete the table. Next, determine the theoretical yield of mass of Ag. From your *actual* mass of Ag and the theoretical yield, determine the % yield.

3. If your ratio $\frac{moles\ Ag\ \text{produced}}{moles\ Cu\ \text{consumed}}$ is greater than the accepted value, then either the moles of Ag is too high, or the moles of Cu is too low. If your ratio is lower than the accepted value, then either the moles of Ag is too low, or the moles of Cu is too high. Use the % yield to help you decide whether the problem is due to the value you have obtained for the moles of silver. List at least one specific experimental error that could account for ratio being too high or too low.

4. It is possible that your value of the ratio $\frac{moles\ Ag}{moles\ AgNO_3}$ is less than 1.00 even if your ratio $\frac{moles\ Ag\ \text{produced}}{moles\ Cu\ \text{consumed}}$ is acceptable. Assuming that you did not lose silver during decanting, what could account for a ratio being smaller than one?

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Stoichiometry Worksheet 2: Percent Yield

For each of the problems below:

a. Write the balanced chemical equation
b. Identify the given (with units) and what you want to find (with units)
c. Show set up with units. Check sig figs, give final answer with units and label.

1. Using the Hoffman apparatus for electrolysis, a chemist decomposes 36 g of water into its gaseous elements. How many grams of hydrogen gas should she get (theoretical yield)?

Bal. Equation:

Before

Change

After

2. Recall that liquid sodium reacts with chlorine gas to produce sodium chloride. You want to produce 581 g of sodium chloride. How many grams of sodium are needed?

3. You eat 180.0 g of glucose (90 M&Ms). If glucose, $C_6H_{12}O_6$, reacts with oxygen gas to produce carbon dioxide and water, how many grams of oxygen will you have to breathe in to burn the glucose?

4. Suppose 4.61 g of zinc was allowed to react with hydrochloric acid to produce zinc chloride and hydrogen gas. How much zinc chloride should you get?
   Suppose that you actually recovered 8.56 g of zinc chloride. What is your percent yield?

5. Determine the mass of carbon dioxide that should be produced in the reaction between 3.74 g of carbon and excess O_2. What is the % yield if 11.34 g of CO_2 is recovered?

6. In the reaction between excess $K_{(s)}$ and 4.28 g of $O_{2(g)}$, potassium oxide is formed . What mass would you *expect* to form (theoretical yield)? If 17.36 g of K_2O is *actually* produced, what is the percent yield?

7. Determine the mass of carbon dioxide one could expect to form (and the percent yield) for the reaction between excess CH_4 and 11.6 g of O_2 if 5.38 g of carbon dioxide gas is produced along with some water vapor.

8. Determine the mass of water vapor you would expect to form (and the percent yield) in the reaction between 15.8 g of NH_3 and excess oxygen to produce water and nitric oxide (NO). The mass of water actually formed is 21.8 g.

Check your answers (1) 4.0 g H_2 (2) 228 g Na (3) 192 g O_2 (4) 9.61 g Zn, 89.1%
(5) 13.73 g CO_2, 82.6% (6) 25.3 g K_2O, 68.8% (7) 7.99 CO_2, 67.4% (8) 25.1 g H_2O, 86.9%

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Chemistry Unit 7
Worksheet 3: Adjusting to Reality - Limiting Reactant

1. Write the balanced equation for the reaction between hydrogen and oxygen.

    Balanced Equation: ______________________________________

    Suppose that 4 molecules of hydrogen gas and 4 molecules of oxygen gas react to form water.

    Make a drawing that represents the reaction container before and after the reaction.

    Before $\longrightarrow$ *After*

____________ How many molecules of water can be produced?
____________ Which reactant is in excess? Why?
____________ How many molecules of excess reactant are there?

Construct a Before-Change-After Table for this reactant mixture:

Bal. Equation: ____

Before: ____

Change: ____

After: ____

According to the table you just made,

____________ How many molecules of water can be produced?
____________ Which reactant is in excess? Why?
____________ How many molecules of excess reactant are there?

Based on your two methods of analysis above, what determines how much product can be made from a particular reactant mix?

2. Write the equation for the formation of ammonia from nitrogen gas and hydrogen gas.

Balanced Equation: ______________________________________

Given 6 molecules of nitrogen and 12 molecules of hydrogen, make a drawing that represents the reaction container before and after the reaction.

Before After

_____________ How many molecules of ammonia can be produced?

_____________ Which reactant is in excess? Why?

___________ How many molecules of excess reactant are there?

Construct a Before-Change-After Table for this reactant mixture:

Bal. Equation: ____

Before: ____

Change:____

After: ____

According to the table you just made,

_____________ How many molecules of ammonia can be produced?
_____________ Which reactant is in excess? Why?
_____________ How many molecules of excess reactant are there?

Describe what you must look for in a particular reactant mixture to decide which reactant will be in excess (have some left over after the reaction):

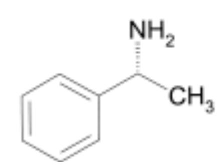
3. When 0.50 mole of aluminum reacts with 0.72 mole of iodine to form aluminum iodide, how many moles of the excess reactant will remain? ____________
How many moles of aluminum iodide will be formed? _____________

Bal. Equation: ____
Before: ____
Change: ____
After: ____

4. When sodium hydroxide reacts with sulfuric acid (H_2SO_4), water and sodium sulfate are the products. Calculate the mass of sodium sulfate produced when 15.5 g of sodium hydroxide are reacted with 46.7 g of sulfuric acid. [Hint: which unit is used in all stoichiometry reasoning?]

5. A 14.6 g sample of oxygen gas is placed in a sealed container with 2.5 g of hydrogen gas. The mixture is sparked, producing water vapor. Calculate the mass of water formed. Calculate the number of moles of the excess reactant remaining.

6. Neuroscientists believe that the only chemical in chocolate that may have a feel-good effect on the human brain is phenylethylamine (PEA). Although the PEA in chocolate occurs naturally, PEA can be made in the laboratory by the following reaction:



phenylethylamine

CH_5NO_2 (ammonium formate) + C_8H_8O (acetophenone) → $C_8H_{11}N$ (phenylethylamine (PEA)) + CO_2 + H_2O

How much PEA can be made from 75.0g of ammonium formate and 125g of acetophenone? What mass of the excess reactant remains?

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4. 27.5 g   5. 16.4 g, 0.34 moles xs   6. 126 g, 9.45g xs

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Chemistry Unit 7 Worksheet 4

Samples of Every Kind of Problem

On a separate sheet of paper, write a complete solution to each of the problems below. Follow the procedure outlined in class. Be sure to circle your final answer.

1. Calculate the number of moles of potassium chlorate, $KClO_3$ (s), that must decompose to produce potassium chloride, KCl (s), and 1.8 moles of oxygen gas.

2. In a single displacement reaction, magnesium metal reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas. How many moles of hydrochloric acid are needed to completely react with 2.43 g of magnesium?

3. Ethane, C_2H_6 reacts with oxygen gas to produce carbon dioxide gas and water vapor. What mass of oxygen gas is required to react with 2.20 moles of ethane?

4. Determine the mass of sodium nitrate produced when 0.73 g of nickel (II) nitrate reacts with sodium hydroxide according to the following equation:
   $Ni(NO_3)_2 + 2\ NaOH \dashrightarrow Ni(OH)_2 + 2\ NaNO_3$

5. In the copper–silver nitrate lab copper metal and silver nitrate solution reacted to produce silver metal and copper(II) nitrate in solution.
   A student placed a copper wire with a mass of 2.93 g in the reaction test tube. The silver nitrate solution contained 1.41 g of silver nitrate.
   He obtained 0.87 g of silver metal. Calculate the percent yield of silver.

6. When hydrochloric acid (HCl) is added to sodium hydrogen carbonate, the products are water, aqueous sodium chloride and carbon dioxide gas. What is the per cent yield if 4.68 g of CO_2 are collected when 10.0 g of sodium hydrogen carbonate reacts with excess HCl?

7. Phosphorus and bromine react vigorously together to form phosphorus tribromide. If 5.0 g of phosphorus and 35 g of bromine react, how many grams of PBr_3 could be produced?

8. Zinc sulfide and oxygen gas react to form zinc oxide and sulfur dioxide. Determine the amount of ZnO that should be produced in a reaction between 46.5 g of ZnS and 13.3 g of oxygen. What is the mass of the xs reactant?

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1. 1.2 moles $KClO_3$ 2. 0.200 moles HCl 3. 246 g O_2 4. 0.68 g $NaNO_3$

5. 97% 6. 89.3% 7. 39 g PBr_3 8. 22.5 g ZnO, 19.5g ZnS xs